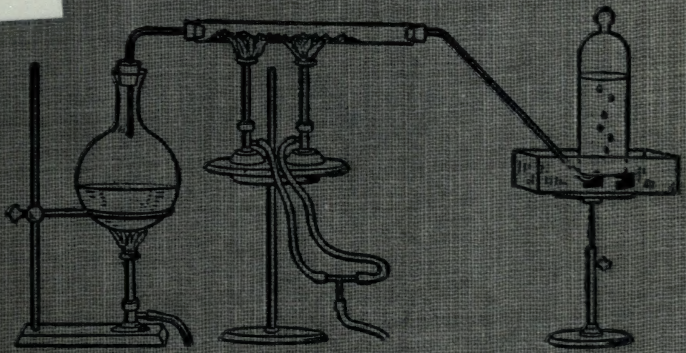
INTRODUCTORY CHEMISTRY

W. S. ELLIS



THE·COPP·CLARK·CO·LIMITED
CANADIAN TEXT BOOKS
TORONTO

INTRODUCTORY CHEMISTRY

SUITABLE FOR USE IN

LOWER SCHOOLS

AND

CONTINUATION CLASSES

BY

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INTRODUCTION.

Chemistry has been prescribed as part of the elementary science work for the lower forms of high schools and for continuation classes. The intention seems to be that the subject should be used at this stage for strictly educational ends rather than for the acquisition of masses of chemical facts and theories by the pupils. If this is correct the objects to be attained by the study will chiefly be facility in manipulation, and training in observation and in accuracy of deduction.

This book has been prepared as an aid in carrying on this work, but only as an aid. There are functions in education that no book can perform, but that are dependent on the personal relations of teacher and pupil. The individuality of the learner will often determine what quantity and what kind of guidance, suggestion and discussion are needed; but direction and supervision are essential elements of any scheme of observational work.

If this course of elementary science is to serve as a borderland between the nature study of the lower school and the more systematic science of later years, it is very necessary that the work be done in such a way as to

develop that scientific spirit of impartial enquiry which is as necessary in literature and history as in chemistry and physics. Experiments in themselves do not accomplish this; indeed, they may be very useless as school exercises, unless care is taken to direct the students' attention to the points to be observed, and to drive the lesson home by discussion and questions.

Lastly, pupils should learn, from the start, that chemistry as a school study derives its chief value from the relationship which it has with the industrial and domestic operations of the period in which we live. If it does not form a connection between the school and what lies beyond the school, it has little claim to a place on the curriculum.

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INTRODUCTORY CHEMISTRY.

CHAPTER I.

SOME PROPERTIES OF GASES.

EXPERIMENT 1.—Fit a large test tube (t. t.) with a stopper, insert a delivery tube bent at right angles, put the mouth of this under water, then heat the t. t. gradually. What takes place at the outlet of the delivery tube?

This result must be due either to a decrease in the volume of the t. t., or an increase in the volume of the gas in it. Which conclusion is the more likely?

Cease applying heat to the t. t., but still keep the mouth of the delivery tube under water. What happens?

Does this result confirm the previous conclusion?

What caused the change?

EXPERIMENT 2.—A long glass tube closed at one end has its open end lowered into water to a depth of 15 to 18 inches. Place a finger carefully over the open end to stop it, then lift the tube out of the water. Did any air escape from the tube? Repeat the operation, but lift the tube slowly out without stopping the end.

In the first case either the tube must have become larger or the air in it must have become smaller in bulk. Which is more likely? What made the change?

EXPERIMENT 3.—Put a little solution of ammonia in a t. t., fit it with stopper and delivery tube, invert another t. t. over the delivery tube, put a strip of moist red litmus paper in the mouth of the inverted tube, then heat the liquid. When the litmus turns a distinct blue, lift off the upper tube and put its mouth under water. Take the stopper out of the other tube and quickly invert it over water.


FIG. 1.

Why does the water rise in the tubes?

The gas that was in the tubes was ammonia mixed with some air.

Point out two differences between ammonia and air.

CHAPTER II.

PHYSICAL AND CHEMICAL CHANGE.

Boil a little water and when steam comes off freely hold a cold plate in it. What forms on the plate? Why?

EXPLANATION.

The water passed through a **physical** change; it became steam, which is still water in a gaseous form.

A lump of iron when heated enough becomes red in color, then white hot; it changes to a more or less pasty condition and may even become liquid, but it never ceases to be iron. The changes are **physical**.

Mention two other examples of physical change among ordinary occurrences at home.

Dissolve about half an ounce of lead acetate (sugar of lead) in a pint of boiling water. Divide the solution into three parts. Put one on a plate and let it evaporate. Put the second in a beaker, lay a splinter across the top of the beaker and suspend a bit of zinc by a thread so that it will be immersed in the solution. Let this stand for twenty-four hours. To the third portion add some iodide of potash solution. The last two cases illustrate **chemical** change, the first physical.

Mention two cases of chemical change in domestic operations at home.

Give two examples of chemical change in manufacturing work.

EXPLANATION.

A substance changes chemically when its constitution is altered so that it becomes a different substance. It may change some of its properties, yet continue to be the same material throughout, then the change is physical. Chemical change is due to addition or separation of matter. Physical change has reference to some properties of the object; chemical change to its composition.

Litmus is a vegetable dye prepared from the root of a plant. It is turned blue by substances that are **alkaline,** and red by those that are **acid.** These two classes of substances, when in contact, tend to neutralize each other, that is, to destroy each other's characteristic properties. Ammonia, soda, potash are common alkalis.

A solid, either in the form of powder or of a mass that is thrown down in a liquid is called a **precipitate.** This word is frequently contracted to **p'p'te.**

Examples of p'p'tes may be obtained by (1) adding drop by drop solution of silver nitrate to a solution of common salt; (2) adding drop by drop sulphuric acid to barium nitrate solution; (3) adding drop by drop bichloride of mercury solution to a little iodide of potash solution.

A substance that produces a decided and characteristic result when applied to another one is said to be a **test** for that other one. Starch paste is a test for free iodine. Litmus is a test to distinguish acids from alkalis. Nitrate of silver is a test for a group of substances to which common salt belongs.

Exercises.

1. Dissolve a little iodine in some alcohol. Place a little more of the iodine in a tube and heat it. Scrape off some of the sediment that forms on the upper part of the tube and try if alcohol will dissolve it. Add a drop of each solution to a little boiled starch. Next take some of the first solution and add caustic potash solution to it until all color disappears, then add a drop of the mixture to some more of the starch. Pure iodine turns starch paste blue. What change did the heating produce in the iodine? What did the potash do?

2. Dissolve a little lead acetate in boiling water. To part of it add, drop by drop, sulphuric acid; to another part add some solution of bichromate of potash. Would nitric acid or hydrochloric produce the same effect as sulphuric? Whether is the p'p'te due simply to an acid or to a particular acid?

3. Make a solution of common table salt. Half fill a t. t. with it and add a drop of nitrate of silver solution. Save this. Half fill another t. t. with the salt solution, add half as much ammonia solution, then drop in silver nitrate. Add a little ammonia to the first tube and shake. Can you account for the different results when the silver nitrate was added?

4. Hold a bit of magnesium wire in the flame. What occurs? What is left? What kind of change went on?

CHAPTER III.

STUDY OF OXYGEN.

EXPERIMENT 1.—Place about half an inch of powdered nitrate of potash (saltpetre) in a small, hard glass t. t.; heat until the nitrate boils, then hold in the mouth of the tube a glowing splinter of wood, such as is obtained by putting the charred end of a half-burnt match in the flame until it rekindles without blazing.

How is the combustion of the match affected? Is there any other gas than ordinary air in the tube?

Repeat the experiment, but when the liquid is boiling freely, drop a small piece of charcoal into it. Hold a similar piece of charcoal in a flame. Compare the combustion in the two cases.

EXPLANATION.

Nitrate of potash when sufficiently heated decomposes, that is, breaks up into simpler substances; one of these is the gas **oxygen.** The burning of a bit of

charcoal, as that on the charred end of a match, indicates the presence of oxygen. If the glowing cinder bursts into flame the gas is mostly, or entirely, oxygen. A bit of glowing charcoal, therefore, serves as a test for this gas.

EXPERIMENT 2.—Try if chlorate of potash may be substituted for nitrate in Experiment 1.

Could a person tell by watching the process, or by the gas given off, whether he was working with chlorate or nitrate?

EXPERIMENT 3.—Heat a little red lead in a hard glass tube. What changes of color does it undergo? Is oxygen given off? How does the material left differ from that taken?

EXPERIMENT 4.—Seal up the end of a bit of glass tubing about as thick as a lead pencil and four inches long. Put into this without smearing the sides a little red oxide of mercury.* Heat it. Watch carefully any changes of color throughout the operation. Is oxygen given off? Examine the grey ring in the tube with a magnifying glass.

Find out the substance forming the ring by holding the tube horizontally, rubbing off the grey matter with a splinter, then inverting the tube over clean paper.

* To introduce a powder into a tube so that there will be no smearing, cut a strip of paper one quarter inch wide and somewhat longer than the tube, fold it lengthwise so that it will form a V-shaped trough; place the powder in this creased paper near one end; carefully thrust this into the tube held horizontally; then turn the whole until the tube is vertical and withdraw the paper.

The material taken was oxide of mercury. What has been obtained from it?

EXPERIMENT 5.—To prepare the gas in quantity proceed as follows:—Set up apparatus shown in Fig. 2;

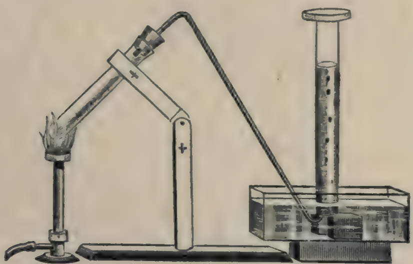

FIG. 2.

put into the t. t. about an ounce of chlorate of potash well powdered and mixed with half its own weight of manganese dioxide; heat strongly.* Collect several tubes full of the gas.

EXPERIMENT 6.—Set one tube of the gas mouth downward over water for several hours. Is the gas soluble?

EXPERIMENT 7.—Thrust a burning splinter into one of the tubes. Does the gas burn? Does the splinter burn?

What change occurs when the splinter passes from air into oxygen?

* When a proper metal retort is available it should be used instead of the test tube; a soup-plate makes quite a satisfactory trough or tank. and wide-mouthed bottles or preserve jars do for collecting vessels. Self-sealers may be used to store the gas for several days, provided the rings and covers fit properly.

EXPERIMENT 8.—Make a chalk cup out of a bit of crayon and a piece of soft wire, as in Fig. 3. Put a shaving of phosphorus in the cup,* ignite it and, while still burning, lower it into a jar of the gas. When the flame dies out, at once turn the jar mouth downward in water and let it stand for some time.

What was the most noticeable thing in connection with the experiment? What becomes of the white fumes? When they disappear hold a piece of paper tightly over the mouth of the jar and turn it up so as to retain the water that has risen in it. Test this water with litmus. What result? What conclusion?

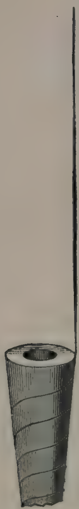

FIG. 3.

EXPERIMENT 9.—The last experiment may be varied by using in the chalk cup powdered sulphur, a small coil of magnesium wire, or a piece of charcoal.

EXPERIMENT 10.—Fill a bottle, of at least 8 oz. capacity, with oxygen. Place a little sulphur on a bit of tin, then heat the end of a piece of braided picture-wire and dip it in the sulphur. Ignite the sulphur that sticks to the wire and hold it well down into the oxygen. What becomes of the wire?

Examine the black substance that dropped to the bottom. Is it iron like the wire? Compare it with the black scale "burnt iron" that is found around a blacksmith's anvil.

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* Keep phosphorous under water, cut it under water and *do not handle it*. Lift it with forceps or on a knife-blade. The heat of the hand is sufficient to ignite it, and its burns are painful and dangerous.

Is there any evidence of great heat having been developed during the action?

What seems to be the most unusual thing that has occurred during this experiment?

What purpose did the sulphur serve in this experiment?

Give any illustration of a similar device in common use.

OXIDATION.

Many substances when properly treated in presence of oxygen enter into combination with it, forming compounds that are called **oxides.**

The process of combination that results in the union of other substances with oxygen is known as **oxidation.** Chemical union that produces light and a large quantity of heat is named **combustion** or **burning.** When substances burn the chemical change is such that an increase of weight may be looked for on account of the oxygen taken up in the combustion.

All ordinary combustion is oxidation, but all oxidation is not combustion.

EXAMPLES OF OXIDES AND OXIDATION.

EXPERIMENT I.—Wet the inside of a bottle that will hold about 12 ozs.; pour out all excess of water and shake clean iron filings about the interior until it is pretty well covered; then stand the bottle mouth downward over water for a couple of days.

Is there any change in the water level? Is there any change apparent in the iron?

If a piece of paper be passed below the bottle and held tightly against the mouth, the bottle may be

inverted without allowing the water that is in it to escape or fresh air to enter it.

Do this and pass a blazing splinter or a lighted candle into the bottle.

Examine carefully the iron filings or whatever remains in their place. What is the color? What does the substance resemble? What happens when a piece of iron is exposed to moist air for some time?

EXPERIMENT 2.—Put about an inch of water in a large bottle, then half a dozen tacks into it, and suspend another half dozen in a bit of muslin so that the cloth will just touch the water. After two or three days examine the two lots.

Repeat the experiment, but weigh the tacks in the bag before they are put in the bottle, and after they have been taken out and dried. What conclusion about the weights of iron and the oxide formed from it?

EXPERIMENT 3.—A glass tube about eight inches long, open at both ends, is filled for nearly half its length with a loose plug of asbestus fibre, a small piece of phosphorus is rolled in asbestus and dropped into the tube, then the other end is closed with a corresponding plug. The whole is carefully balanced on a scale; then the tube is heated near the phosphorus; after combustion has ceased and the tube cooled it is again weighed. The white smoke should be all retained by the asbestus.

How is the weight affected by the burning?

Why is it necessary to have the tube open at both ends?

If the asbestus had not been used what would have become of the oxide produced?

Compare the burning of a shovelful of coal.

EXPLANATION.

About one fifth of the bulk of the atmosphere is made up of oxygen gas. Iron is one of the substances that combine with oxygen at ordinary temperatures, especially if moisture be present. In Experiment 1 the water rose because the oxygen, or most of it, left the air to form a reddish solid with the iron. This solid is commonly called rust, but is chemically one of the oxides of iron. It occurs in great beds in many parts of the earth, and is then called haematite (or hematite), a very valuable iron ore.

When the iron-wire burned in oxygen it formed black oxide or magnetic oxide of iron mostly. This substance also occurs very extensively as a mineral, and is mined largely as another ore of iron.

Exercises.

1. Pour a little limewater into a bottle of oxygen and shake it with the gas. Lower a bit of burning charcoal, by means of a wire, into the bottle, put a piece of moist paper on the mouth to prevent the gas escaping. After combustion has ceased, again shake the water with the gas. Why should the result be different from the former one? Was there any difference in the appearance of the gas in the bottle? Was there any actual difference? What became of the charcoal?

2. Dip some tacks in oil, then expose them to moist air, as was done before. What is the result?

Why should metal-work and implements that are exposed to the weather be painted? Machinery that is to lie idle for a time is generally coated with oil, vaseline, or some such substance. Why?

3. The breaking down of wood during rotting is partly a process of oxidation in presence of moisture. Account for the custom of painting wooden structures.

4. Mention some common examples of oxidation (*a*) in domestic operations; (*b*) in industrial work.

5. What devices are you acquainted with for bringing the oxygen of the air into contact with burning substances?

6. What advantages does man derive from the combustibility of coal, wood, and oil?

7. Dissolve a shaving of phosphorus in carbon disulphide, spill the solution over a piece of filter paper and watch it while the liquid evaporates. Why is this result different from that obtained when a shaving of phosphorus was put on the paper?

CHAPTER IV.

STUDY OF AIR.

EXPERIMENT I.—Set up apparatus as in Fig. 4. Fill the flask with cold water that has been freely exposed to the air for several hours. The tube should not go quite through the stopper and the water should stand just above the stopper in the tube. Gently heat the flask. When gas ceases to pass over, invert the receiver, without allowing air to enter it. Test the gas with a glowing splinter and with a bit of burning phosphorus in a chalk cup.

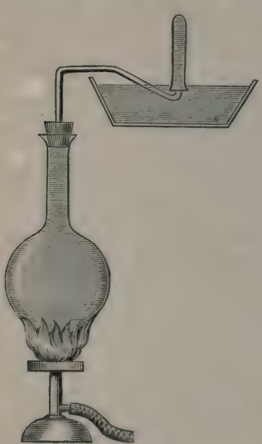

FIG. 4.

Does the gas resemble air in its appearance and in its action with burning substances?

Try the experiment with water that has recently been boiled.

Assuming that the gas is air, where did it come from? "Fishes will die without air." How do they obtain their supply of it? Fish in tanks must have frequent changes of water. Why?

EXPERIMENT 2.—Make a stand about three inches high and one inch across out of wire, as in Fig. 5.

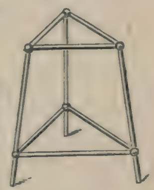

FIG. 5.

Set this in a dish that has some water in it. Lay a bit of tin on top of the stand; put a little phosphorus on the tin; ignite it, and turn a wide-mouthed bottle over the whole, so that the mouth of the bottle is under water.

Let it stand until combustion ceases and the white fumes disappear.

Does the gas that remains look like air?

Invert the bottle without allowing air to enter it.

Test the gas with a glowing splinter. See if any phosphorus remained unburned.

The white fumes in this case were the same as those formed when phosphorus was burned in oxygen.

Why did the phosphorus cease to burn? Why does the splinter not burn? Why did the water rise in the bottle?

EXPERIMENT 3.—Use the appliances of the last experiment, but instead of phosphorus put some wet iron filings on the tin. Let the whole stand for a couple of days. Again test the gas that remains with a burning splinter or candle.

What is the water that passes into the inverted bottle the measure of?

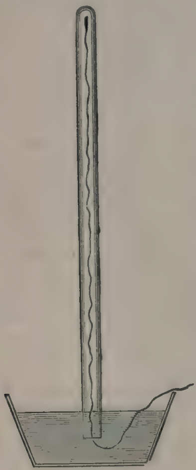

FIG. 6.

EXPERIMENT 4.—A tube about twenty inches long, closed at one end, as in Fig. 6, has a piece of freshly-cut phosphorus passed to the closed end by a wire, and the open end is at once put under water. This is allowed to stand for forty-eight hours; then the open end of the tube is immersed in water, without being removed from the vessel A, until the water is at the same level within the tube and without it.

Measure the height the water has risen in the tube. From this calculate what percentage of the gas has disappeared.

Why is it necessary to get the water level the same within and without the tube?

EXPLANATION.

The atmosphere is mainly composed of a mixture of two gases, **oxygen** and **nitrogen.** The oxygen in the experiments disappeared, as gas, in the formation of the oxides of iron and phosphorus. The nitrogen remained as it does not easily enter into combinations. Besides these two gases, air contains small portions of other substances, particularly **water vapor** and an **oxide of carbon.**

USES OF OXYGEN.

All ordinary combustion is oxidation; the oxygen being supplied by the air and the product being an

oxide of the substance burned. Many of the operations of ordinary life could not be carried on if it were not for the supply of oxygen in the atmosphere. Most of the heating of our homes, and the lighting of our houses and streets, as well as the generating of power for transportation and manufacturing, are all possible because this gas is a constituent of the air.

All animal and most vegetable life is dependent on a supply of oxygen. In special organs, lungs or gills, the oxygen acts upon the blood and keeps it in a proper condition for performing its functions in the body. During this process the oxygen becomes combined with carbon to form carbon dioxide, the same gas that is produced when charcoal burns, and the one that turns limewater white. Plants also absorb oxygen for use in their growth processes.

Exercises.

Twelve parts by weight of carbon unite with thirty-two parts of oxygen. Assuming that coal is 80 % carbon, how many pounds of oxygen will be required to burn a ton of it ?

Explain the use of the front damper of a stove. Why is it placed below the fire ?

Why is the base of the burner of a coal-oil lamp punched full of holes ?

Give instances of domestic and industrial operations that are dependent on oxidation.

Mix a little powdered chlorate of potash and charcoal ; ignite the mixture.

Account for the vigorous combustion.

Heat some chlorate of potash in a tube until it boils, then drop in the burnt end of a match.

CHAPTER V.

STUDY OF HYDROGEN.

EXPERIMENT 1.—Put into a t. t. some water and add to it about one quarter its volume of sulphuric acid. What change in the temperature of the water?

Drop some bits of zinc into the tube. What happens?

Cover the mouth of the tube with a piece of wet paper for a minute, then remove the paper and quickly touch the mouth of the tube to a flame.

Is the escaping gas air? Is it oxygen?

What reason for the answers to these questions?

What becomes of the zinc?

Will zinc and water produce a gas?

Will acid and water?

Will zinc and undiluted acid?

EXPERIMENT 2.—Fit up apparatus as in Fig. 7. Put into the large tube zinc and dilute sulphuric acid. After all air has been expelled collect several tubes full of the gas.

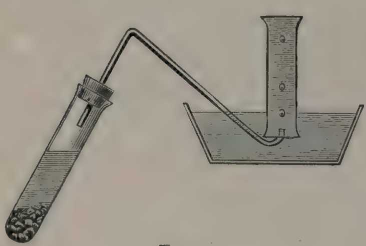

FIG. 7.

Hold two of the tubes side by side, one mouth upward, one mouth downward, for two or three minutes, then bring the mouth of each to a flame.

Note—When sulphuric acid is to be mixed with water it should always be done by pouring the acid slowly into the water, not by adding water to the acid.

How does hydrogen compare in weight with air?

Is the gas perceptibly soluble in water?

Pour some of the liquid remaining from the preparation of hydrogen into a beaker and let it stand over night. Evaporate a little of this liquid to dryness.

EXPLANATION.

In the preparation of hydrogen the water dissolves this white substance which otherwise would coat the pieces of zinc and prevent the acid from coming in contact with the metal.

OTHER METHODS OF PREPARING HYDROGEN.

EXPERIMENT 1.—Drop some pieces of iron wire into hydrochloric acid, heat the vessel and collect the escaping gas over water. Test it.

EXPERIMENT 2.—Fill a dinner plate with water; drop a piece of sodium on it. Float a bit of red litmus-paper on the water and put a small bit of sodium on it. What condition is the sodium in when the lump becomes globular?

EXPLANATION.

Sodium combines with water to form an alkaline solution and hydrogen gas is given off. The sodium is difficult to control on account of its fusibility and its tendency to dance about on the water. To overcome this the sodium is made into an amalgam which

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Note—Hydrogen forms a dangerously explosive mixture with air. If the gas is to be ignited it should be tested before a flame is brought near any of it in a closed vessel. This testing is done by collecting a small t. t. full of the gas and holding it to a flame. If it burns with a sharp explosion or with a whistling report it is mixed with air. If it ignites almost noiselessly a flame may safely be applied to a quantity of the gas.

resembles a solution of sodium in mercury, except that it is solid when in certain proportions. The water acts with the sodium of the amalgam quite irrespective of the mercury.

The movements of the lump of sodium on the water are due to the formation of hydrogen whenever the metal touches the water. The gas, when produced, throws the lump of sodium to one side; immediately more hydrogen is formed below it and the metal is tossed to another place. The action is similar to that which occurs when a drop of water falls on a hot stove-lid, except that in the latter case steam, not hydrogen, is the propelling gas.

If the sodium be held in one place, as when retained by the fibres of the paper, it will apparently burst into flame. It is really the hydrogen gas which burns and the yellow color is given to it by some sodium vapor mixed with it. When the globule of metal is kept in one spot the chemical action between it and the water develops heat enough to cause the hydrogen to ignite, but when the sodium is moving about there is not a high enough temperature at any one time to cause the union of the escaping hydrogen with the oxygen of the air.

Exercises.

Will some zinc dropped into hydrochloric acid give hydrogen?

Try magnesium and sulphuric acid.

Will nitric acid and zinc give hydrogen?

Make some sodium amalgam by heating some mercury to boiling in a tube, then dropping in small bits of sodium, one at a time. Turn the fluid out on a cold plate. If sodium enough has been

added the substance will be a brittle grey metallic looking mass. Put some pieces of it in water and collect the escaping gas in a tube. Test it for hydrogen. What remains from the amalgam?

Test the water with litmus paper.

Roll a bit of sodium, as big as half a pea, in sheet lead, such as is used to line tea chests; make some pin holes in the lead, and drop it into water. Collect the gas that rises. Test it.

CHAPTER VI.

ATOMS AND MOLECULES.

A theory universally held by chemists and physicists is that any form of matter, whether solid, liquid or gaseous, is composed of indefinitely minute portions, each one of which has the properties peculiar to the substance of which it is a part. The exact conditions under which these infinitesimal portions exist is but ill understood even by the great scientists, but for the purposes of this book it will be sufficient to imagine them as infinitely small particles, each existing by itself and endowed with powers of attraction and repulsion for other particles. These smallest parts in the case of an element are called **atoms.**

ELEMENTS AND COMPOUNDS.

Substances that do not consist, so far as is known, of two or more different kinds of matter in union are named **elements.** Those that result from the union of two or more different kinds of matter are called **compounds.** No one has ever succeeded in getting from iron anything but iron except by causing it to

combine with something else. Iron cannot be decomposed into other substances. It is, therefore, an element. So is hydrogen, oxygen, magnesium and sodium. About seventy-five elements are known.

A substance that is formed by the union of other simpler ones is a **compound.** When magnesium burned it formed a white ash by the combination of the metal with oxygen. When iron was exposed to damp air it rusted, that is, combined with oxygen. These are examples of compounds.

Some of the more commonly-occurring elements are oxygen and nitrogen as gases of the air; carbon, as coal and graphite; sulphur, and some of the metals that are found native as copper, silver, gold. Examples of well-known compounds are water, all common rocks, all organic substances, and the gases formed by burning fuels.

Atoms have the power of forming themselves into groups of definite numbers, and every such group is called a **molecule.** In rare cases, as will be learned later, a molecule may be monatomic, that is, made up of one atom; but in elementary work the molecule may be taken as a definite group. Thus two atoms of hydrogen, when free to act combine with one, and only one, atom of oxygen. The white fumes formed when phosphorus burned were composed of molecules made up of two atoms of phosphorus joined with five of oxygen.

All chemical action is a re-arrangement of the atoms composing the molecules of the substances affected.

This will be illustrated after chemical notation has been explained.

CHAPTER VII.

CHEMICAL SYMBOLS AND FORMULAS.

In arithmetic it would be practically impossible to write down in words all numbers used and all operations performed; so a system of symbols, called figures and signs, has been invented to represent briefly the quantities and the operations. Similarly, in chemistry, symbols are employed to represent substances and the chemical reactions among them.

The following are the names of a few of the more common elements and the symbols that stand for them:

Carbon	C.	Oxygen	O.
Chlorine	Cl.	Phosphorus	P.
Copper	Cu.	Potassium	K.
Hydrogen	H.	Silver	Ag.
Lead	Pb.	Sodium	Na.
Magnesium	Mg.	Sulphur	S.
Mercury	Hg.	Zinc	Zn.
Nitrogen	N.		

It will be noticed that generally the symbol consists of the initial letter of the name. When, however, a number of names begin with the same letter, as carbon, chlorine, cerium, caesium, cobalt, calcium, cadmium, chromium, it would be clearly impossible to indicate all by the same letter, C. In such cases the oldest and best known element of the group generally has the initial letter for its symbol, and the others are indicated by two letters, consisting of the initial and the one next prominent in the sound of the word. Thus, in the group above the symbols are C, Cl, Ce, Cs, Co, Ca, Cd, Cr. In such cases

the first letter is a capital, the other a small one. This does not account, however, for such odd symbols as those of silver and sodium. The explanation of these is that they are formed from the old Latin names of the substances. Thus, in the list

Cu (copper) comes from the Latin cuprum.

Pb (lead) comes from the Latin plumbum.

Hg (mercury) comes from the Latin hydrargyrum.

K (potassium) comes from the Latin kalium.

Ag (silver) comes from the Latin argentum.

Na (sodium) comes from the Latin natrium.

The signs " + " and " = " are also of very general use, but have a significance somewhat different from that which they have in mathematics. The former may be translated by the words mixed with, or in contact with; the latter by the words yield, or produce, or give, or form.

Quantitatively the symbol H means one atom of hydrogen; C, one atom of carbon; 2O, two atoms of oxygen; 3Cl, three atoms of chlorine. H_2 means a molecule consisting of two atoms of hydrogen; Cl_2, a molecule consisting of two atoms of chlorine.

The difference between such symbols as 2H and H_2 may be illustrated by magnetized needles stuck in bits of cork and floating in water. A pair of them drifting about separately, without any connection, would be two individual things, and would represent the atomic condition; but if this pair, oppositely magnetized, came together, and floated as a group of two, that would represent the molecular state.

A group of symbols written side by side, without signs, means that the atoms represented by those symbols are in chemical union. The group is called a formula, and stands for one molecule of the compound formed. H_2SO_4 means one molecule of a compound, and that molecule is composed of two atoms of hydrogen, one of sulphur, and four of oxygen.

$$H_2SO_4 + Zn = ZnSO_4 + 2H.$$

One molecule of H_2SO_4 with one atom of zinc gives one molecule of $ZnSO_4$ and two atoms of hydrogen.

Some illustrations may now be given that chemical action is a re-arrangement of atoms to form different molecular groupings.

When oxide of mercury was heated the attachment between the oxygen and the mercury atoms was broken, and each kind of matter went off by itself. This is indicated by the formulas $HgO = Hg + O$.

A molecule consisting of one atom of mercury and one of oxygen is disrupted, and the two atoms separate. In the experiment this happened with a great number of the molecules, so, many atoms of oxygen and of mercury were set free.

This kind of action is called a **decomposition.**

In the preparation of hydrogen the chemical action is represented by the equation

$$Zn + H_2SO_4 = ZnSO_4 + 2H.$$

Here the zinc and hydrogen change places, and finally the SO_4 is combined with Zn and the two atoms of hydrogen are set free.

CHAPTER VIII.

STUDY OF WATER.

EXPERIMENT 1.—By means of a perforated stopper and rubber tubing attach a burner to a hydrogen generating apparatus. The burner may be a common blow pipe, the inner tube of an oxyhydrogen jet, or a bit of glass tubing drawn out fine. When it is safe to do so (page 17) ignite the escaping hydrogen and hold a dry cold bottle mouth downward over the flame for some time.

Is smoke given off from burning hydrogen?

Is anything given off?

On what evidence is the answer based?

When hydrogen burns what chemical action goes on?

What is the product of the combustion?

Mention cases in which water exists as an invisible gas.

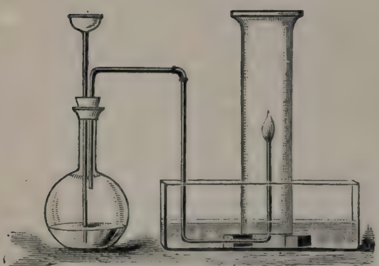

FIG. 8.

This experiment illustrates the preparation of water by causing the elements that form it to combine. These are hydrogen and oxygen.

EXPERIMENT 2.—Arrange apparatus as in Fig. 8. When the hydrogen jet is lit the jar at the right is inverted over it, full of air. At the instant the hydrogen ceases to burn take the stopper out of the flask, so that no more

gas will pass into the jar. Invert the jar without letting air into it. Test the gas with a burning splinter. What is the conclusion?

Why did the hydrogen jet go out? What does the rise of the water measure?

A current of hydrogen constantly passed through the tube from the flask. Why did the volume of gas in the jar diminish then?

EXPERIMENT 3.—Connect a four-cell battery in series. Bring the ends of the leading-out wires nearly together in some water that has a little sulphuric acid in it. What change takes place?

Connect the wires to a decomposition of water apparatus as in Fig. 9. (If this or some similar piece is not

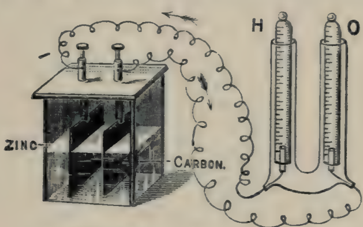

FIG. 9.

available, flatten out the ends of the wires and pass them under two test tubes filled with acidulated water and held mouth downward close together in a vessel of the same.) Do equal quantities of gas collect in the two tubes?

How much greater is one volume than the other?

Test the gases. What are they? Which one came off in greater quantity? Were either of these the same gas that came off in Experiment 1, page 16?

EXPLANATION.

The first experiment illustrated the composition of water by bringing its elements together and causing them to combine; this one shows the decomposition of water through the agency of the electric current. In the former case hydrogen and oxygen united to form water, at first as an invisible vapor; in the latter one water is divided into its constituent gases, oxygen and hydrogen.

Water is not easily decomposed, hence its oxygen is not readily available for oxidation purposes. Phosphorus that quickly combines with oxygen of the air does not join with that of water at all; and iron under water rusts but slightly while in air it rusts rapidly.

EXPERIMENT 4.—Arrange apparatus as in Fig. 10.

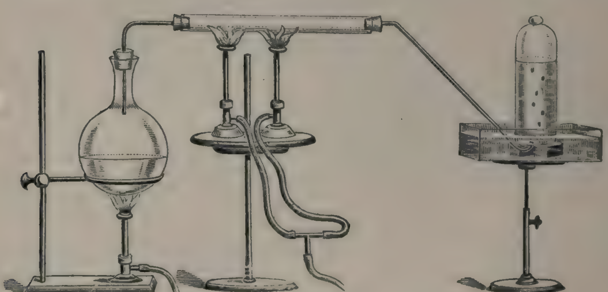

FIG. 10.

In the horizontal tube put some fine iron filings; heat them to redness; at the same time pass steam from the flask of boiling water over them.

Examine the gas that passes over. Test it.

When the experiment is done turn out the contents of the large tube; compare these with some of the original filings.

EXPLANATION.

Some metals that at ordinary temperatures do not decompose water will do so, if heated to redness or a higher degree. Iron is one of these. This metal forms two common oxides, red oxide or rust, and black oxide. The latter, in a somewhat impure form, may be picked up about a blacksmith's anvil as black scale or "burnt iron." In machine shops where steam hammers are used, much larger plates of this scale may be found. The term burnt iron is scientifically correct. The iron when white hot undergoes oxidation in the air. In the experiment the hot iron decomposed the water passing over it, took up the oxygen and allowed the hydrogen to pass on. Some of the steam in hot iron pipes is similarly broken up into its constituents.

In the experiment the iron is oxidized, while the water is **de-oxidized** or **reduced.**

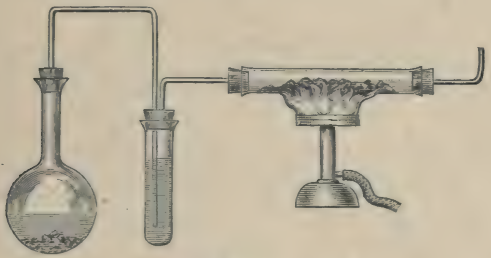

FIG. 11.

EXPERIMENT 5.—In the apparatus shown in Fig. 11 the hydrogen generated in the left-hand flask is led through strong sulphuric acid slowly to dry it, then it is passed through a tube containing black oxide of copper; when this is heated to redness hold a cold, dry glass bottle over the outlet.

What change in the copper oxide?

What passed off?

What chemical action went on?

Was there oxidation?

Was there reduction?

Why should the hydrogen be dried?

What remained in the tube?

Heat a little of this red substance in air.

EXPERIMENT 6.—Repeat the last experiment, but use powdered and well-dried iron rust instead of the copper oxide.

How is the change of color accounted for?

Trace the oxidation and de-oxidation that go on.

Would the heating alone produce the alteration in the rust? Try it.

Will hydrogen alone reduce copper oxide to copper? What property does hydrogen possess that enables it, by the aid of heat, to break up copper oxide?

EXPERIMENT 7.—Place some red lead in the tube instead of the iron rust. First heat the lead to redness, let it cool and note any apparent change. Then pass hydrogen over it and again heat to redness.

EXPLANATION.

There are three common oxides of lead, viz.:—the yellow oxide, PbO, the red oxide, Pb_3O_4, and the brown oxide, PbO_2. When brown oxide or red oxide is heated, the chemical changes are indicated by the following equations respectively:—

$$PbO_2 = PbO + O.$$

$$Pb_3O_4 = 3PbO + O.$$

In the experiment what was the effect of the hydrogen? What gas passed off when the red lead was heated?

What was given off as gas when the hydrogen was introduced?

USES OF WATER.

Water is an essential material for animal existence It is requisite for digestion and for supplying the fluids lost through surface evaporation and breathing. The water held in a gaseous form in the atmosphere contributes to man's comfort by checking excessive evaporation from the skin and mucous membrane.

Water is the most general solvent for solids, and thus contributes to plant life by entering into the vegetable tissues and carrying with it dissolved mineral matters needed for vigorous growth. All soils, except in very limited arid regions, are full of moisture that is constantly acting on the soluble solids and preparing them for plant consumption.

Closely connected with this solvent power is the washing process going on in every rainstorm which cleanses the atmosphere of suspended impurities and carries enormous quantities of decaying vegetable matters down into the soil. The quantity of water that is lifted by evaporation and carried by the air currents to other regions where it is precipitated as rain or snow is so great that it is difficult to express it in a way that will be comprehended. For instance, it does not require a very heavy day's rain to give a fall of half an inch; that is, if no water ran away or soaked into the ground there would be a coating one half inch thick.

Such a storm, however, would mean a fall of fifty-six tons of water on every acre. Two feet of rain, therefore, during the year over the whole of Canada is a quantity unthinkable on account of its greatness; yet year by year this enormous mass of water is transferred by the heat of the sun and by winds from oceans and lakes, from ponds and soil, to upland districts, there to nourish forests and crops, and gradually to find its way back to the lower levels, turning mills and carrying commerce as it goes.

The wearing down of rocks into soils and the distribution of the latter are also largely due to water action. Every ice crystal that forms in a rocky crevice tends to break up the stone by expansion; and the washings of the hills are distributed over the lower levels at every flood.

The employment of water in domestic operations is within the experience of everyone. Its chief value here lies in its solvent power, in its absorption by solids through capillarity, and in its having a boiling point that is constant for any degree of heat above 212° Fahr. or 100° C.

Its value in industrial operations, and generally as a working agent, arises from the comparatively low temperature at which it takes the gaseous form. Its abundance and the ease with which it may be thus transformed make the steam engine a commercial possibility.

Exercises.

Point out the part which water plays in (1) the existence and growth of plants; (2) the lives of animals; (3) modifications of climate; (4) communication between different parts of the earth.

What advantages come to man (1) from the power of water to dissolve some of the solids; (2) from the small range of temperature through which water passes in going from the solid to the gaseous condition?

A piece of iron waterpipe, one inch in diameter and about a foot long, may be used to show the explosive effect of a mixture of oxygen and hydrogen. Close one end of the pipe with a common cork, not a rubber one. Then fill the pipe with a mixture of gases, two parts hydrogen, one part oxygen (a less violent result will be obtained if a mixture of one part hydrogen and three parts air be taken). Close the other end of the pipe with a cork that has a hole bored in it to serve as a vent. Hold the vent to a flame, but be sure that the pipe is in such position that if the corks fly out they will not do injury.

EXPLANATION.

That which is commonly called an explosion is generally only a case of very rapid burning, but this is often accompanied by the formation of a large volume of hot gases. Gunpowder, for instance, is a mixture of substances that burns with great violence, and gives off a large quantity of gaseous products. In the case of oxygen and hydrogen the explosion is due to the rapid heating of the gases to a high temperature at the moment that the flame runs through them.

CHAPTER IX.

ACIDS, BASES, SALTS.

The term **acid** generally suggests sourness, probably because the common acids are all sour to the taste. They also change blue litmus solution to red, but the chemist regards both of these properties as not essential to the existence of an acid. In his view an acid is a

compound, each molecule of which contains one or more atoms of hydrogen that may be displaced by atoms of metals when properly treated. As an example, zinc displaced the hydrogen of sulphuric acid, thus :—

$$Zn + H_2SO_4 = ZnSO_4 + 2H.$$

The formulas for some common acids are HNO_3, nitric acid; H_2SO_4, sulphuric acid; HCl, hydrochloric acid, H_2CO_3 carbonic acid, H_2SO_3 sulphurous acid.

When the hydrogen of these acids is replaced by potassium the resulting compounds are

$$KNO_3,\ K_2SO_4,\ KCl,\ K_2CO_3 \text{ and } K_2SO_3.$$

These are **salts.** The term **salt** is given to those substances formed when the hydrogen of an acid is replaced by a metal. The displacing metal is called a **base.** The acid radical, SO_4, NO_3, SO_3, etc., forms part of the salt molecule, but is then joined with one or more atoms of some metal instead of hydrogen.

In naming salts the suffix, **ic,** of the acid name is changed to **ate** in that of the salt; thus, from nitric acid is formed nitrates, from sulphuric acid sulphates, from acetic acid, acetates, from carbonic acid carbonates, and so on through the list.

Those acids whose names end in **ous** form salts whose names end in **ite**; thus, sulphurous acid gives sulphites, nitrous acid, nitrites.

If a substance consists of but two elements its name ends in **ide.**

The following are some illustrations of names of substances :—Na_2SO_4, sulphate of sodium or sodium sulphate; $FeSO_4$, sulphate of iron; $FeCO_3$, iron carbonate;

KNO_3, potassium nitrate; KNO_2, potassium nitrite; $MgSO_4$ magnesium sulphate; $MgSO_3$ magnesium sulphite; $MgCl_2$ magnesium chloride; H_2O, hydrogen oxide; CaS, calcium sulphide.

In some chemical names the prefixes **mon**, **di** or **bi** and **tri** are introduced to indicate respectively one, two or three parts; thus, CO and CO_2, are carbon monoxide and carbon dioxide; SO_2 and SO_3 are sulphur dioxide and sulphur trioxide; $FeCl_2$, $FeCl_3$ are dichloride, and trichloride of iron respectively.

The following are the chemical names of a few common substances:—Salt is sodium chloride, NaCl; washing soda is sodium carbonate, Na_2Co_3; baking soda is sodium bicarbonate, $NaHCo_3$; sal-ammoniac is ammonium chloride, NH_4Cl; marble and limestone are calcium carbonate, $CaCO_3$; lime is calcium oxide, CaO, when fresh, but calcium hydroxide, $Ca(HO)_2$, when slaked; copperas is iron sulphate, $FeSO_4$; saltpetre is potassium nitrate, KNO_3; and plaster of Paris is calcium sulphate, $CaSO_4$.

CHAPTER X.

STUDY OF CARBON.

EXPERIMENT I.—Shake up a little limewater with the air in a bottle, then wrap the end of a piece of wire round a lump of charcoal and hold it in a flame until it is burning freely; lower the charcoal into the bottle. After a short time withdraw the charcoal and again shake the gas in the bottle with limewater.

What evidence is there of the gases in the bottle having been changed?

The lump of charcoal grew smaller while burning. What became of the part that disappeared?

Explanation.

When carbon burns it forms a compound with the oxygen of the air, CO_2, known by the names carbon dioxide or carbonic-acid gas. Limewater is used as a test for the presence of this gas, as the two together form a white, insoluble compound, usually described as *milky*, and having the composition $CaCO_3$. (This is the precipitated chalk of the druggist.)

Experiment 2.—Test for carbon dioxide the gas that comes out of the top of a lamp chimney. By means of a bit of tubing blow through limewater for a couple of minutes. Hold an inverted t. t. above a gas flame or alcohol flame for a few moments; test the contents of t. t. with limewater. What do the results show about the composition of the breath and of the gas from the chimney?

Whence was the carbon derived in the case of the lamp?

When a lamp *burns* what chemical change goes on?

What purposes do the wick, the chimney, the oil, and the dome of the burner respectively serve?

Is there carbon in coal oil, in alcohol, in coal gas?

Experiment 3.—Put a little washing soda (baking soda will do just as well) in a t. t. Pour a little very dilute acid on it and, by means of a cork and delivery tube, lead the gas that comes off into limewater, What is the gas?

Does soda contain carbon? Does it contain oxygen?

On what evidence is the answer to these questions based?

Try if a bit of limestone will yield the same gas when treated with acid.

From these experiments devise an easy method for preparing carbon dioxide in quantity.

From observations thus far, if one wanted a steady stream of the gas for some time, what materials should he choose?

EXPERIMENT 4.—Fill a large t. t. with the gas, invert it over water and let it stand for a couple of days.

Test the gas with litmus. To do this, place two bits of wet litmus paper, one blue, one red, in a t. t. filled with the gas, then invert the tube over water and let it stand for a time.

What conclusion about the solubility of the gas?

Test the water that rises in the t. t. with litmus.

EXPERIMENT 5.—Fill a jar with the gas, lower a lighted candle into it. What conclusion?

Repeat the experiment, but use a blazing splinter instead of the candle.

Does the gas burn?

Does it permit substances like a candle or a splinter to burn in it?

EXPERIMENT 6.—Suspend a burning candle, by means of a wire, near the bottom of a tall vessel, as a wide-mouthed bottle. Fill another vessel of equal size with carbon dioxide gas, then try if the gas can be poured

like water into the vessel containing the candle. Tilt the vessel full of gas slowly.

A better form of this experiment is the following:— Pour a little limewater into a bottle, then bring the open end of a delivery tube that is carrying a current of carbon dioxide just over the mouth of the bottle. Notice if a milky layer forms at the top of the limewater after a time.

What information do these experiments convey about the relative weights of air and carbon dioxide?

EXPERIMENT 7.—Fill a large bottle with water, carry it into the open air, pour the water out, put about a tablespoonful of limewater in the bottle and shake it well. Turn the limewater out into a t. t. Does it show any whitening?

Repeat the experiment except that the bottle is to be emptied in a badly-ventilated room in which the air has become *close* or foul. Does the limewater show whitening? Why fill the bottle with water?

Connect this result with that of Experiment 2.

Mention one purpose that ventilation serves.

Mention two sources of vitiated air in common living rooms.

Let some limewater stand in an open vessel without shaking it, in a class-room, for a couple of days. Examine the crust that forms on it. Put a drop of dilute acid on some bits of this crust.

What happens?

What is the crust, and how was it formed?

Does limewater when exposed to the air outside the building form a similar crust?

Explain the results.

Write out particulars regarding carbon dioxide under the headings, state at ordinary temperatures, appearance, solubility, combustibility, power to support combustion, weight as compared with air.

How could this gas be distinguished from hydrogen, from oxygen, from nitrogen, from air?

EXPLANATION.

When carbon burns in a free and plentiful supply of oxygen or air it forms carbon dioxide, CO_2; but if the quantity of oxygen is limited, carbon monoxide, CO, is produced.

Carbon dioxide may be partially de-oxidized by passing over white-hot carbon, thus:—$CO_2 + C = 2CO$. This is a reduction from a higher oxide to a lower one. On the other hand, the monoxide will combine with oxygen and burn to CO_2, thus undergoing oxidation again. For instance, the flickering blue flame that plays over the surface of a coal fire when fresh fuel is put on is carbon monoxide burning to the dioxide.

When carbon dioxide dissolves in water it forms carbonic acid, H_2CO_3. The salts of this acid constitute a very large part of the earth's crust. Thus calcium carbonate is found as limestone, marble, chalk, coral and animal shells. Magnesium carbonate forms very large bodies of rock in some parts of the world. Carbonate of iron (siderite) is an important ore of that

metal while the carbonates of sodium are the washing soda and baking soda of domestic operations.

Any carbonate when treated with one of the strong acids will effervesce, giving off carbon dioxide; sometimes heat must be applied. **This is the common test for a carbonate.** Compare Experiment 3.

Exercises.

1. Pass a current of CO_2 into limewater until it forms a thick white p'p'te. Let this settle, pour off the clear liquid then add to the solid a little dilute acid. The following equations indicate the chemical actions, if hydrochloric acid is used. Express them in words.

[$Ca(OH)_2$ is calcium hydroxide, slaked lime, which dissolves in water.]

$$Ca(OH)_2 + CO_2 + H_2O = Ca(OH)_2 + H_2CO_3$$
$$= CaCO_3 + 2H_2O.$$
$$CaCO_3 + 2HCl = CaCl_2 + H_2CO_3$$
$$= CaCl_2 + H_2O + CO_2.$$

2. Will strong hydrochloric acid affect oystershell, clamshell, snailshell, eggshell?

3. Pass a current of CO_2 into limewater for some time after the white p'p'te has formed.

How is the clearing of the liquid accounted for?

What has become of the white solid?

Is the liquid acid?

Add a few drops of an alkali, as ammonia or caustic potash.

What is the effect of this alkali on the acid?

Why does the white p'p'te reappear?

USES AND APPLICATION OF CARBON DIOXIDE.

Whenever wood, coal, coal oil, or illuminating gas is burned in air carbon dioxide is produced. A ton of coal, for instance, sets free three and two-third tons of

this gas; and, as an excess of it is injurious to animal life, it is possible that the atmosphere might become vitiated by this means if provision was not made for the elimination of the carbon dioxide. This provision is found in the vegetable world, whose supply of food is nearly all derived from this gas in the atmosphere. During summer time when plant growth takes place every green leaf that is spread in the air absorbs carbon dioxide which is decomposed through the agency of the chlorophyll, the green particles in the leaf cells; the oxygen is mostly exhaled while the carbon goes to build up the plant tissues. Plants and animals, therefore, largely counteract each others' effect on the atmosphere.

When a bottle of soda water, is unstoppered there is a rapid outflow of gas that often carries some of the fluid with it as foam. The gas in this case is carbon dioxide that was dissolved in the liquid under great pressure. When the stopper is removed the pressure is relieved and the gas escapes from solution. This is typical of all effervescing drinks. In most of them the gas is artificially prepared; in wines and beers, however, its presence is due to fermentation, a process of decomposition that results from the action of plants such as yeast.

In preparing bread, carbon dioxide is set free in the dough, and the expansion of the mass due to the formation of the gas within it causes the bread to *rise.* The "lightness" of the bread depends on the even distribution of the bubbles of gas and the hardening of the dough at the proper time by baking. The carbon dioxide may come from a carbonate contained in a baking powder or from the growth of immense numbers of yeast plants that have been sown within the mass.

FORMS AND OCCURRENCE OF CARBON.

Wood roasted, that is heated to a high temperature out of contact with air, yields charcoal which is nearly pure carbon. The black cinder often left when a burning stick goes out is an example.

Coal is a mineral carbon, somewhat impure. Coke is nearly pure carbon, obtained by driving off some of the ingredients of coal. Graphite or plumbago, sometimes called blacklead, is another mineral carbon.

Diamonds are masses of crystalline carbon.

Charcoal does not occur in nature, but has to be prepared artificially when wanted.

Coal exists in all stages of formation, from lignites that are only one step removed from peaty masses, through bituminous and cannel coals to anthracite. Coal beds are found in certain rock formations, generally limestone, and exist not as veins breaking through the rock beds, but as sheets spread out between the rock layers. Graphite is found in scattered masses, usually small scales, amid the granite formations. It is used largely for polishing purposes and as a lubricant for heavy machinery. The "lead" of lead pencils is graphite generally ground up, made into a paste, and moulded into the form in which it is mounted in the wooden case.

EXPERIMENT I.—Half-fill a t. t. with pine cuttings; place half an inch of chalk in small lumps above them, then heat the tube for about fifteen minutes to redness.

Test the escaping gases with litmus.

Try if the gas will burn.

Does it contain CO_2?

After heating to redness for some time, spill out the contents of the tube.

What purpose did the chalk serve?

What is left of the wood?

Is it still wood?

What changes are noticeable?

Will the residue burn?

Will it burn up entirely?

Is charcoal pure carbon?

What has been learned about the volatile constituents of wood?

What about the non-volatile part?

Are these statements correct in the case of hardwood—maple, for instance?

EXPERIMENT 2.—In the apparatus, Fig. 12, the left-hand tube contains some pieces of soft coal damped with chalk as before; the other one is full of water to wash the escaping gas. Heat the coal strongly for about fifteen minutes.

Test the gas that comes off with litmus.

Is the gas inflammable?

If the escaping gas will burn, hold a t. t., mouth downwards, over the flame, then shake the contents with limewater.

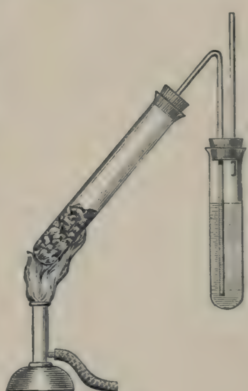

FIG. 12.

Judging from its odor, what is the brown substance that collects in the wash water.

Drain off a little of the clear wash water, or filter some of it out, then add a few drops of caustic potash and heat the mixture. Can any escaping gas be detected by its odor.

Apply to some of the wash water the test given on page 50.

Examine the substance left in the tube that was heated.

Will it burn now?

Explain why soft coal burns with a flame when put upon a grate fire. Will the brown, sticky substance burn? When coal is burned in a grate or stove why does this brown substance not appear?

EXPLANATION.

When bituminous coal is roasted a number of volatile substances are driven off. Some of these condense at ordinary temperatures, some are soluble, some are combustible. The first group are mostly tarry products, the second are largely ammonia compounds, and the third is a mixture of substances that make up illuminating gas. The non-volatile part of the coal is coke. In gasworks this operation of roasting coal is carried on in large retorts for the object of securing the gaseous substances that, after purification, are burned in houses for light and heat. Coal tar and salts of ammonia are by-products, and the coke is a valuable fuel.

SOME USES OF COAL.

The chief uses of coal arise from its property of giving off large quantities of heat when undergoing chemical union with oxygen. To this is due its value as a fuel.

Those kinds of coal that have only a small portion of volatile substances in their composition burn with but little flame ; but those that give off much gas burn with luminous and often smoky flames.

When coal or charcoal is heated in contact with substances that contain much oxygen the coal will frequently combine with the oxygen, thus breaking up the other compound. This is the basis of smelting operations. Ores that are either oxides, or that are easily changed into oxides, are then mixed with broken coal and heated strongly. The coal combines with the oxygen generally to form carbon monoxide and the metal is set free.

A moment's consideration will make it clear that the progress of mankind is dependent on coal more than on any other single substance except, perhaps, water. The carrying on of domestic operations, the possibility of large population existing in northern climates in winter, the conducting of great manufacturing industries, the carrying of the world's products to suitable markets, either by land or sea, are all possible because of the simple chemical fact, that carbon combines readily with oxygen and gives off a large quantity of heat in the act of union. Almost the entire power supply of the world to-day is due to (1) the property of coal just mentioned, (2) that water is easily changed into the gaseous condition, and (3) that gases are expansible and elastic. These account for the use of steam as a motive power. Another great source of energy arises from the ready movement of portions of water under the influence of any external force. Thus with the mobility of water and the attraction of gravitation masses of water at high

levels can easily be made to do work. Here, again, the supply at high levels is due to easy vaporization and to atmospheric drift.

Flame.

Luminous flames for these experiments may be obtained either from an ordinary fish-tail gas burner, a Bunsen burner with the holes closed, a coal-oil lamp or a spirit lamp in which some turpentine or dissolved camphor is mixed with the alcohol. A non-luminous flame may be had by using a Bunsen burner with the holes open or an alcohol lamp with methylated spirits only.

Experiment 1.—Hold a lump of clean, cold metal in a non-luminous flame for a minute.

Examine the surface of the metal. Did any black smoke come off from the flame when the metal was thrust into it?

Repeat the experiment with a luminous flame.

What is the result this time?

Experiment 2.—Hold a splinter of wood across the non-luminous flame of a gas burner about half an inch above the tube. An alcohol flame may be used if the top of the wick is spread out a little and the splinter held a quarter of an inch above it.

Does the splinter char equally at all parts within the flame?

Where is combustion most rapid?

Repeat the experiment with luminous flames.

Do the results differ?

Repeat again, but instead of the splinter of wood use a sheet of white paper thrust horizontally across the flame, for a few seconds, at the same height as before.

If this is successful there should be a charred ring on the paper, but the middle part unburned.

Repeat the experiment but pass splinter and paper across the flame halfway between its top and bottom.

EXPERIMENT 3.—Bring a piece of fine wire gauze down on top of a flame. Turn on the gas; hold a piece of fine wire gauze about an inch above the burner, then try to light the gas on the upper side of the gauze.

How are the results explained?

Read a description of the structure and principle of the safety lamp used by coalminers.

EXPLANATION.

Flame is produced by burning gas. The gases that are commonly burned are all compounds containing carbon as one of their constituents.

The substances composing these gases do not burn with equal vigor, so that if the supply of oxygen be limited, or if the burning mass be cooled below the temperature at which all the constituents are consumed, one or more of them may escape unburned.

Coal gas or the vapor of coal oil is a mixture of substances each consisting of hydrogen and carbon (hydrocarbons).

In the case of a luminous flame the carbon particles are apparently separated from the hydrogen and are heated up to whiteness in the body of the flame, thus giving out light; hence the flame is luminous. These

carbon particles are finally burned up when they come in contact with the air in the flickering yellowish margin of the upper part of the flame. When the mass of burning gas is cooled, as by the lump of metal, the carbon escapes unburned into the air in the form of smoke or soot. If too little oxygen be supplied to the burning gas, carbon will also escape.

Alcohol yields a non-luminous flame, or nearly so, because the vapor contains relatively little carbon. Turpentine, benzine and camphor are rich in this element; consequently, if any of these be mixed with the alcohol the flame will become luminous.

If oxygen, sufficient for complete combustion, be mixed with a gas the carbon particles will be at once consumed without being made white hot for any appreciable time; so the flame is non-luminous. This accounts for the blue gas flame of a Bunsen burner when the holes are open.

Exercises.

A coal-oil lamp smokes when the flame is turned too high. Why?

What is the smoke in this case?

When the chimney is taken off a lamp it will smoke. For what reason?

If a lamp be carried from a warm room to a very cold one the flame will apparently shrink. Why?

How is the flame of a lamp affected by putting a piece of paper on the top of the chimney?

How, by holding a piece of paper against the bottom of the burner?

Account for the result observed in each case.

The term lampblack is given to carbon deposited on cold surfaces. How may it be obtained?

Make three rolls of waste paper, such as is obtained by rolling up a leaf of a scribbler; saturate one with alcohol, one with turpentine, one with coal oil. Set fire to them separately and watch whether the paper ignites at once or whether the liquid burns off first. Is there any perceptible difference between the three flames.

Bring a piece of cold metal gradually up against the side of a candle flame.

Does the flame stay in contact with the metal?

When one blows out a candle, how is the burning stopped?

How is the same result secured by turning down a lamp?

What does the black smoke that is often seen coming out of factory chimneys consist of?

Why does it come out at some times and not at others?

If it were formed from cooled gases containing carbon, what condition in the furnace might produce it?

CHAPTER XI.

STUDY OF AMMONIA.

EXPERIMENT 1—Make a solution of sal-ammoniac, (ammonium chloride), drop into it some caustic potash and heat. What gas is escaping?

Test it with litmus.

EXPLANATION.

Ammonia gas may be prepared by mixing sal-ammoniac and lime and using water enough on the mixture to make a thin paste, then heating.

It is obtained, however, much more easily by heating the ammonia liquor which is a solution of ammonia gas in water that may be obtained at the drug stores.

The origin of the ammonia salts of commerce, as ammonium nitrate, chloride, sulphate, etc., is the soluble constituents of coal gas as it leaves the retorts. The gas liquor, as the wash water is called, is treated for the separation of ammonia products.

EXPERIMENT 2.—Fit a large t. t. with stopper and delivery tube, and put a little ammonia solution in it; heat this, and when the gas is coming off pass it into another tube that has been filled with hydrochloric acid vapor by heating a few drops of the acid in it.

How are the white fumes accounted for?

EXPERIMENT 3.—Put about half an inch of ammonia solution in a large t. t., heat this until all air is expelled then quickly invert the tube, placing the mouth under water.

Why does the water rise in the tube?

How could it be shown that the rising of the water is not due to the cooling of the gas?

Will ammonia gas burn in air?

Try if a blazing splinter will continue to burn when immersed in the gas?

To what extent is the gas soluble?

EXPERIMENT 4.—Take two exactly equal portions of ammonia solution, then dilute some hydrochloric acid, using about three parts of water to one of acid; put this in a t. t. and mark its upper edge with a bit of gummed paper. Drop a bit of litmus paper into one of the ammonia tubes and add the acid very slowly, with constant shaking of the mixture until the litmus is just changing to red.

The acid and alkali are now said to be **neutralized.**

Note how much of the acid was required by measuring from the original height to the present level.

Boil the other portion of ammonia for a couple of minutes. After it has cooled neutralize it in the same way.

Does it require as much of the acid as the former one did?

What conclusion about the strength of the solution after boiling?

What effect has heating on the quantity of gas held in solution?

If a strong solution of the gas had to be prepared which would it be better to use, warm water or cold?

Could a solution of the gas be prepared in water containing acid? Why?

Compare this experiment with Experiment 1, Chapter IV.

Can the results be similarly explained?

EXPERIMENT 5.—Put a little solid sal-ammoniac in a t. t. and heat it.

Does ammonia come off from it?

Does it melt?

Does it vaporize?

Compare ammonia with hydrogen, oxygen, nitrogen, air, carbon dioxide, stating in what ways it resembles each of these, in what respects it differs from each, and how it could be distinguished from any of them.

If some oxygen were suspected of containing ammonia gas, how could the latter be removed, if present?

Ammonia is commonly collected by being passed upwards into an inverted vessel. Suppose a delivery tube were thus carrying the gas into a bottle, how could one tell when the bottle was full?

Why may ammonia not be collected over water, as hydrogen is?

Nessler's test for ammonia is prepared by adding solution of bichloride of mercury (corrosive sublimate), to solution of iodide of potash until the p'p'te just dissolves; then add to this ten times its bulk of strong caustic potash solution.

A yellow or reddish discoloration in a liquid caused by this test indicates the presence of ammonia.

Try how many drops of ammonia solution in a quart of water will show the first trace of discoloration by use of this test.

USES OF AMMONIA.

Since ammonia is strongly alkaline it is a valuable and much-used cleansing agent because of its joining with fatty substances to form soluble compounds.

There is a well-known principle in physics that when a substance changes from one state to another in which the molecules are less rigidly bound together, heat is absorbed, and is again set free when the reverse molecular process occurs. In accordance with this, when a substance alters from a solid to a liquid, or from a liquid (either in fusion or solution) to a gas, heat is absorbed from surrounding substances.

Ammonia is very largely soluble in cold water, and is very volatile; hence it is a particularly-convenient material for employment in cooling operations.

Generally, the plan adopted is to pump ammonia gas from a strong solution in one part of the apparatus through pipes into another part where it is again dissolved under pressure.

The rapid volatilization of the ammonia passing into the gaseous state reduces the temperature of the surrounding material, whether air or water.

The shipment of perishable goods, as dairy products, fresh meats, fruits, etc., for long distances, has been made possible by refrigerating processes in cars and steamers. Cars and storehouses may use ice harvested in winter, but factories, ships, and large warehouses find it more economical to obtain the necessary cooling by artificial means.

In tropical countries ice is readily made though it never occurs in nature.

In large hotels, ocean steamships, and other such places, a refrigerating plant, using ammonia, is a regular part of the outfit.

The compound formed when ammonia is neutralized by hydrochloric acid, and the result evaporated, is sal-ammoniac or ammonium chloride. It is used in some manufacturing works, and is employed in electric batteries of the type used for telephones and bells.

CHAPTER XII.

STUDY OF LIMESTONE.

EXPERIMENT 1.—Dilute some hydrochloric acid one half with water, then drop into it a bit of common limestone.

Collect a couple of test tubes full of the gas.

Does it burn?

Will a splinter burn in it?

Test it with limewater.

What is it?

If any solid matter is left in the tube, filter it out.

Is this solid soluble in acid?

Is it of the same composition as the rest of the rock?

Evaporate a few drops of the clear liquid to dryness.

What remains?

Is limestone soluble in water?

EXPERIMENT 2.—Drop some crystalline limestone (marble or calcite) into dilute hydrochloric acid.

What is left in the tube?

Drop some sulphuric acid into the clear liquid.

Try to dissolve some of the rock in sulphuric acid.

Repeat, using dilute nitric acid.

In what respects do the results with the three acids resemble each other?

In what do they differ?

EXPLANATION.

Limestone is chemically **carbonate** of **calcium** with mixtures of impurities, chiefly clayey substances. In the crystalline form it is much purer, the crystals themselves being nearly free from foreign matter.

When the carbonate is treated with hydrochloric acid, the substances formed are carbon dioxide and calcium chloride ; the latter is a soluble solid.

If sulphuric acid be used the substance formed is calcium sulphate which is a white insoluble solid. The nitrate is soluble.

EXPERIMENT 3.—Try if limestone will affect the color of moist litmus paper. Try also if water in which broken limestone has been left standing for some hours will change the color of litmus.

Wind a piece of wire about a bit of the rock and hold it in the flame until it becomes coated with a grey powder. Let it cool, then scrape a little of the powder on moistened red litmus paper.

Put some more of the powder in a few drops of water and test it with litmus.

Does heating to redness change limestone chemically ?

On what evidence is the answer based ?

Is the substance soluble which is formed during the operation ?

EXPLANATION.

Several of the carbonates when heated undergo decomposition by parting with carbon dioxide. Lime-

stone, calcium carbonate, is one of these, and the chemical action is indicated by the equation $CaCO_3 = CaO + CO_2$.

The CaO, **oxide of calcium,** which is left behind is **quick lime.**

The carbonate is insoluble, but the oxide is slightly soluble in water.

EXPERIMENT 4.—Take a lump of freshly-burned quick lime, or lime that has been kept carefully from exposure to air; drop water on it gradually.

Does the temperature change?

How much water had to be poured on before the lime showed signs of being wet? (Weigh the lime before and after.)

Repeat the experiment but use dried lime that has been wet, or lime that has been exposed to the air for several days.

Is the result different from that in the former case?

EXPLANATION.

Oxide of calcium unites chemically with oxide of hydrogen to form a combined oxide of calcium and hydrogen, thus: $CaO + H_2O = CaH_2O_2$. This formula is frequently written $Ca(HO)_2$ and in this form means that the oxygen of calcium oxide has been replaced by two groups of HO, known as **hydroxyl radicals,** and the compound is named **calcium hydroxide,** or **calcium hydrate.** Commonly it is called slaked lime. When, therefore, water is added to freshly-burned lime, the first change is chemical and results in the production of the hydroxide; the next one is physical because the

hydroxide becomes somewhat dissolved, as may be shown by alkalinity of the water in which lime has been standing.

EXPERIMENT 5.—Drop some lime into dilute hydrochloric acid.

Does the lime disappear?

Does gas come off?

Is the chemical action the same as when limestone is treated with acid?

To find whether the action is a mere dissolving of the lime in water, or is a chemical one, add lime in excess; let the mixture stand for an hour, then try what effect the liquid has on litmus.

EXPERIMENT 6.—Mortar is made of lime, sand and water.

Does fresh lime give off gas when treated with acid?

Does old lime give off a gas with acid?

Does sand give off gas when dropped into acid?

Secure from some old building a piece of mortar that has long been exposed to air. Try if it will yield a gas when dropped into dilute acid.

If it does, determine whether the gas is hydrogen, oxygen or carbon dioxide.

Is old mortar anything different from a mixture of lime, sand and water?

What reasons for the answer?

EXPLANATION.

It is true that fresh or "green" mortar is a mixture of slaked lime, sand and enough water to make it into a paste. The hardening process which this undergoes is, in part at least, a re-formation of the carbonate of calcium by absorption of carbon dioxide from the air. When slaked lime is exposed to carbon dioxide it gradually combines with it; and, as the atmosphere contains this gas, as well as water vapor, the calcium hydrate is constantly exposed to the slight action of carbonic acid, which with the alkaline base forms the carbonate. There are perhaps other chemical changes, but that mentioned is the chief one. The first stage in the "setting" of mortar, that is the hardening of it, is probably only a drying process.

Of late years a kind of lime known as **cement** has come into very general use in construction work, road-making, etc. The material for this cement is a mixture of calcium carbonate and clay. For this supply the marl beds found in many parts of the country are largely worked. **Marl**, when pure, is the broken shells of small aquatic animals, that apparently lived in bodies of water which once occupied the districts where the beds are found. **Clay** is largely a compound of silicon and aluminum, and is mostly ground up and disintegrated felspar, one of the constituents of granitic rocks.

In some cases the marl beds contain clay; in others, the mixture has to be made artificially. The substances are ground, mixed with coal dust, and burned or calcined. The mass is finally ground very fine; and when it is made into a paste with sand and water it has the

property of "setting" somewhat slowly into a very hard and durable mass; hence its value in construction work.

USES OF LIMESTONE AND OTHER FORMS OF CALCIUM CARBONATE.

Limestone is a valuable building material, chiefly because of its durability, the ease with which it may be quarried, and the even thickness of the beds over considerable areas. In this connection crystallized limestone, as marble, is especially prized. The use of marble for pillars, monuments, etc., depends on its property of taking a high polish.

The uses of lime and cement turn upon their setting into a hard solid when made into a paste with water and sand, so that they bind adjoining masses of stone or brick together. Both the decomposition of the limestone and the hardening of the pasty mass are chemical actions.

Lime is used also as a fertilizer for some crops, as calcium is thus supplied in a soluble form to plants that require it.

Closely related to limestone is another calcium compound, the sulphate, which occurs also in bedded layers as gypsum. When calcined it forms plaster of Paris which has the property of quickly "setting" when made into a paste with water.

In this case the rapid hardening arises from the absorption of water in the formation of crystals. The variety of calcium sulphate, which is found in nature in the crystallized form, is selenite.

CHAPTER XIII.

CRYSTALLIZATION.

EXPERIMENT I.—Make a saturated solution of copper sulphate, pour the liquid on a plate and allow the water to evaporate at the ordinary temperature of the room.

What is the most noticeable thing about the solid that remains?

Again prepare a saturated solution, pour the liquid into a beaker and suspend in it by a thread a bit of solid sulphate.

What is the result?

EXPLANATION.

Most inorganic substances, when they take the solid condition either after fusion or solution, tend to arrange their molecules in definite ways so that the resulting mass has a symmetrical geometric form. The thin plates of the copper sulphate were all rhomboidal, for instance. These geometric masses are called **crystals.** The property of taking characteristic shapes in crystallizing is made use of in determining unknown substances, because **each has its own crystalline form.**

In most cases the crystals run together in masses, thus suffering distortion, so that their particular form cannot be easily made out. In the case of solutions the crystals begin to form only when the solution is saturated. If it is dilute, evaporation will go on until the saturation point is reached. Anything that will serve as a nucleus about which the molecules can arrange themselves will in a proper solution become the centre

about which a crystal, or more likely a cluster of crystals, will grow. Substances that have no crystalline form are said to be **amorphous** (formless). Examples are glass, resin, gum, coal, paraffin.

EXPERIMENT 2.—Make a saturated solution of bichromate of potash in boiling water; let the solution stand in a beaker until it cools.

Why should so much of the solid separate out in the cooling?

Examine the ring at the top of the liquid with a magnifying glass. Does it show any crystallization?

When the liquid gets cool pour part of it on a plate and let it evaporate.

Suspend a little bit of the solid salt in some more of the solution.

Compare the form of the crystals on the plate with those obtained from the copper sulphate.

EXPERIMENT 3.—Prepare crystals of common salt, white sugar, permanganate of potash, alum and iodide of potash, by pouring saturated solutions separately on plates.

Try to obtain crystals of alum by suspending a bit of the salt in a saturated solution.

Examine these crystals carefully, using a magnifying glass if necessary.

Make drawings of them.

Note.—Such crystals as these may be easily mounted for preservation by warming a plate of glass, then running a thin coat of melted paraffin over it. Crystals placed on this waxed surface will become embedded if the plate is warmed until the paraffin melts. If the wax coat is thin the name of the substance may be written with a needle. Such mounts should be kept in a closed case in which there is an open vessel of water.

How do the salt crystals differ from those of sugar?

EXPERIMENT 4.—Warm a dry plate of glass, spill over it a hot saturated solution of sal-ammoniac (ammonium chloride); quickly drain off the surplus fluid. A growth of lace-like crystals should take place on the wetted surface. If the glass be laid on a black ground this growth may be readily seen.

EXPERIMENT 5.—Make saturated solutions of copper sulphate, bichromate of potash and alum, mix them and pour the mixture on a plate. After evaporation what is left?

What conclusion should be come to about mixtures of substances and the crystals they form?

If a substance such as copper sulphate were impure and if one were to dissolve it and crystallize the salt how would this affect the purity?

If a person wanted a drug free from all foreign matter in which form, crystalline or powdered, would he be more likely to get it?

WATER OF CRYSTALLIZATION.

EXPERIMENT 6.—Prepare some crystals of copper sulphate; see that they are perfectly dry. Put one about half an inch square or several smaller ones in a dry t. t.; heat this holding the tube nearly horizontally. If water does not form on the cold part of the tube repeat, but put the mouth of this tube into another one a little larger.

What changes come over the crystals in regard to form and color?

Turn the solid out on a plate and let a drop of water fall on it. What change?

Was there water to be seen in or about the crystals to begin with?

Break some of the crystals to be sure that they do not enclose water.

Was there water in the crystals?

On what evidence is the answer based?

Repeat the experiment using alum crystals either whole or broken. What conclusion?

If the crystals were ground to powder would they still contain water?

Again repeat, but use sulphate of magnesium (Epsom salts).

Does this result agree with others?

Try if washing soda acts similarly.

EXPLANATION.

In the formation of many kinds of crystals, especially in the case of salts, water appears to be necessary; thus magnesium sulphate is $MgSO_4 + 7H_2O$, that is, the molecule of sulphate is joined with seven molecules of water when it goes to form crystals; and **if the water is not present the sulphate will not crystallize.** Similarly with a molecule of alum there are twenty-four molecules of water. The water is present in the crystal in **solid form,** and when it is driven off by heating the crystal breaks down to powder.

The water thus taking part in the formation of crystals is known as **water of crystallization.**

Put some crystals of washing soda in an open vessel where they will be exposed to dry air for some days.

The crystals should change into a white powder, thus $Na_2CO_3 + 10H_2O$ loses the ten molecules of water and becomes pure Na_2CO_3.

A substance is said to **effloresce** when the crystals thus fall away to powder through the loss of water of crystallization.

CRYSTALLIZATION OF WATER.

Examine snowflakes on a day when big, feathery ones are falling from low-drifting clouds, with little or no wind blowing.

Examine the frost feathers that appear on windows on very cold days.

Observe the ice lances that form at edges of pools when they first begin to freeze.

Study the white rime that covers the twigs on foggy mornings when the temperature is below freezing.

These are all examples of the crystallization of water when it assumes the solid state, that is, when it passes from the fused condition into the solid by cooling. Ice is probably always made up of such crystals, but on account of crowding their individuality is lost in the mass. Every snowbank is but a heap of broken and crushed water crystals, and a lump of ice is made of a countless number of them pressed into a solid cake by the expansive force exerted by the change of state.

METALLIC CRYSTALS.

Some of the metals take a crystalline condition under proper treatment.

EXPERIMENT 7.—Make a solution of lead acetate in boiling water and hang a bit of zinc in it when it cools.

After twenty-four hours examine the metallic mass without removing it from the water.

Is there evidence of crystallization?

Exercises.

If crystals of minerals and ores can be obtained, study their forms. Suitable ones are quartz, iron pyrites, calcite, fluor spar, tourmaline, micas, hornblendes, apatite, magnetite, galena, selenite.

Why should frost ferns form on the inside of store windows and not on the outside?

Breathe against a cold window pane on a very cold day. The feathery, crystalline mass that should form closely resembles that obtained from the sal-ammoniac on the glass.

FORMATION OF CRYSTALS.

For the formation of perfect crystals, whether from fusion or solution plenty of room is necessary; hence in masses of material the perfect crystals are found generally at the surface.

In order that the crystals may be large the evaporation or cooling must go on slowly.

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